Ch 6 Chemical Bonding

Basis for all structures in life Why things stay together

Chemical bond → link between atoms that results from mutual attraction of their nuclei for e^-

Types of bonds → (nonmetal + metal) (metal + polyatomic ion)

1. **Ionic bonds** - attraction b/twen pos + neg ions resulting in formation of a chemical bond.

e^- are transferred Ex NaCl

2. **Covalent Bonding** - Chemical Bond resulting in sharing of electrons

2 nonmetals (polyatomic ion)

Bonds are seldom completely ionic or covalent

That Beautiful Grey area

The for Ionic bonds we consider E-neg

The larger difference in Electroneg is,

More ionic the bond is

The e⁻ will spend more time around the more e⁻ neg atom

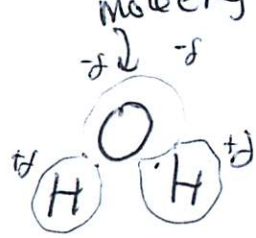
Degrees of covalent bonds

↳ ~~More ionic characters — nonpolar covalent~~

Polar covalent → bond where united atoms ~~are~~ have unequal attraction for e⁻

Polar → means separation of charge.

So some ~~have~~ covalent bonds have slight separation of charge → H_2O



Non polar covalent - bond where e⁻ are shared equally b/tween atoms.

~~Measure~~ Tell by using following chart

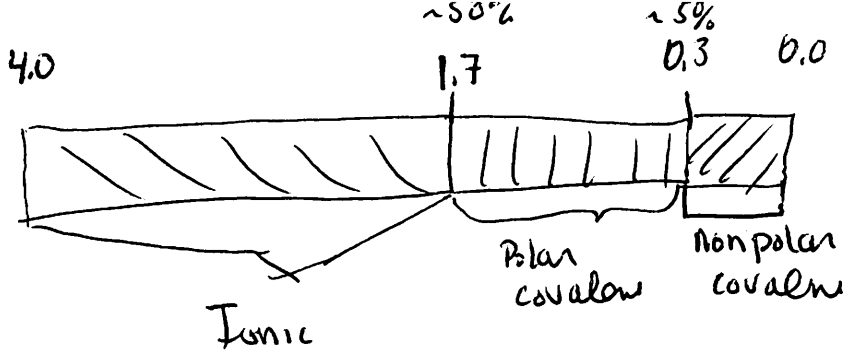


Fig 6-2

To determine types of bond
subtract e-neg and if between

4.0 - 1.8 Ionic

1.7 - 0.3 Polar Covalent

0.3 - 0.0 non polar covalent

Ex Sulfur eneg = 2.5 More e neg →

Type of bond w/ H? (2.1) 2.5 - 2.1 = .4 Polar covalent sulfur

" C_s ? (0.7) = 2.5 - (.7) = 1.8 Ionic sulfur

Cl (3.0) = 3.0 - 2.5 = .5 Polar cov Chl

Between B(2.0) + H(2.1)

2.1 - 2.0 = 0.1 nopolar

13 Variation of Ionic Character Nonpolar and Polar-Covalent Bonds

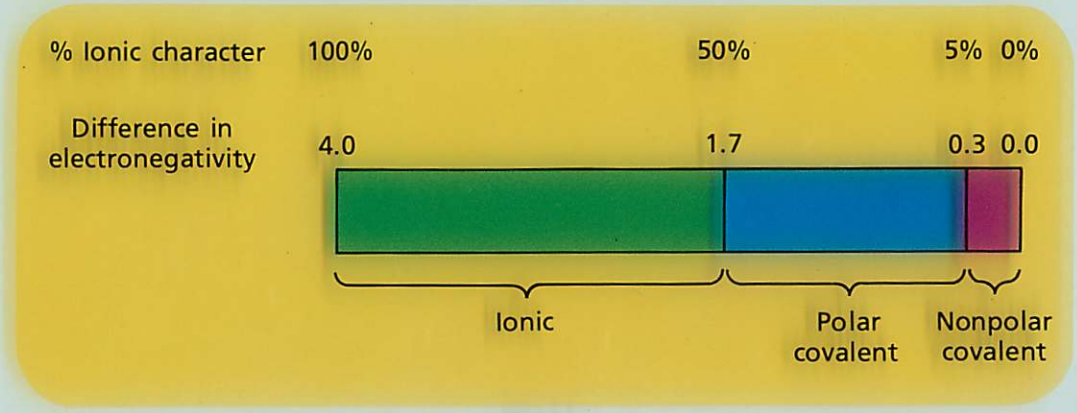


Figure 6-2

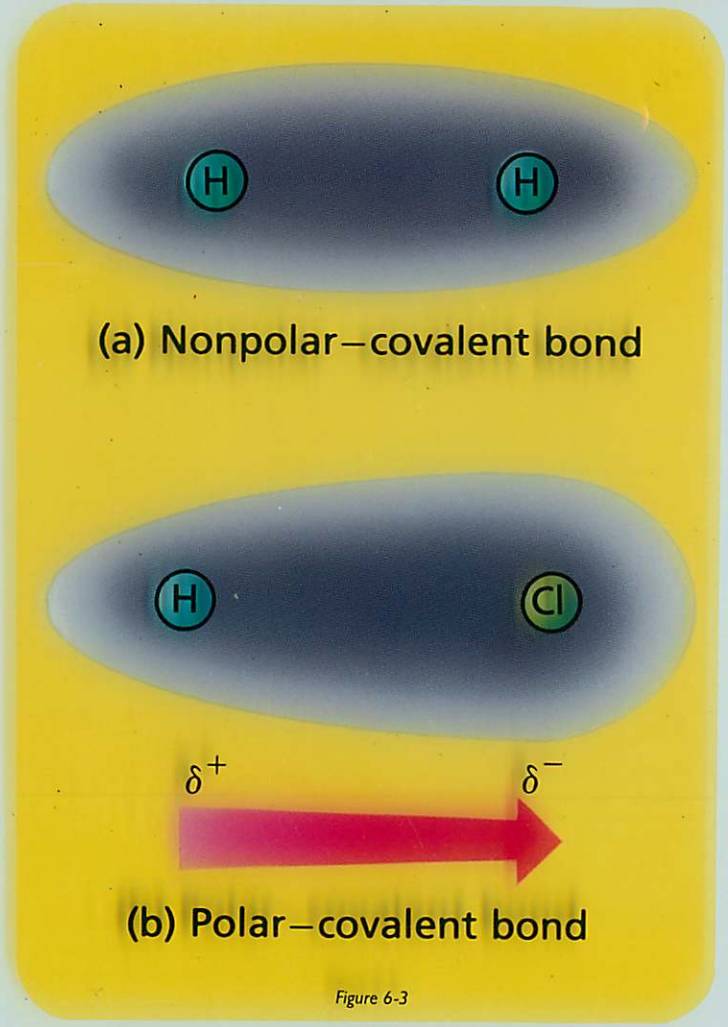
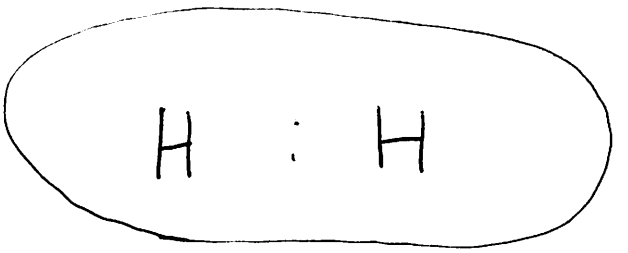
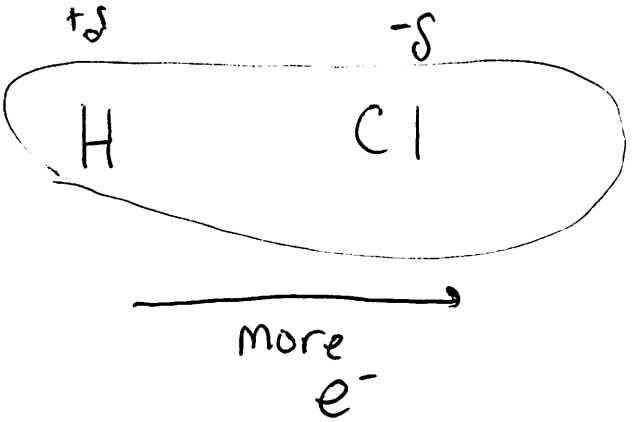


Figure 6-3

Density of e^- between Polar + Nonpolar



Nonpolar
equal distribution



Polar

Why do bonds form?

b/c of decrease in PE

If E given off means bond is more favorable. Any time PE is lowered that state is favored

pg 164 Q2, 3, 4

6.2 Covalent Bonding + Molecular Comps

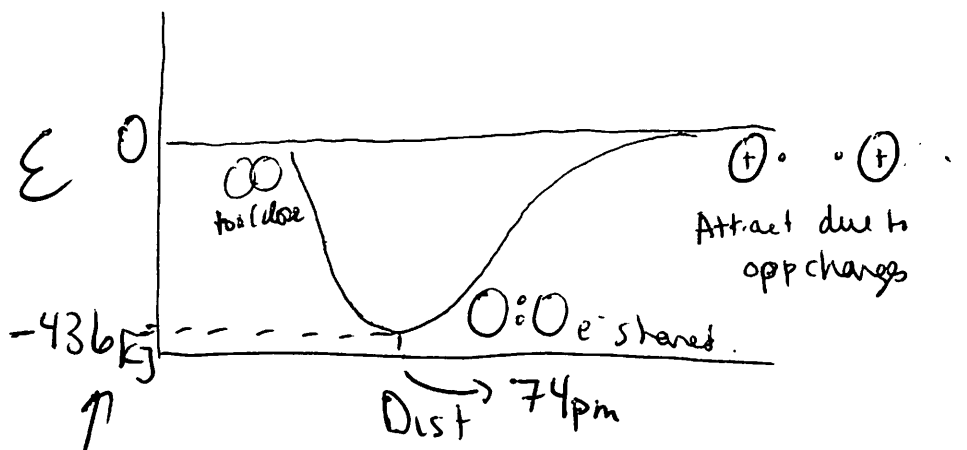
Molecule → group of 2 or more atoms held together by covalent bonds and able to exist independently

diatomic molecule → Compound w/ 2 atoms

Chemical formula - shorthand rep of the composition a substance using atomic symbols + # subscripts

H_2O, H_2, $C_6H_{12}O_6$

How a bond forms? Energy diagram



436 kJ of E released

↳ Release of E lowers PE puts molecule @ lower PE than atoms stable

14 Bond Energy Diagram and Forces on Atoms

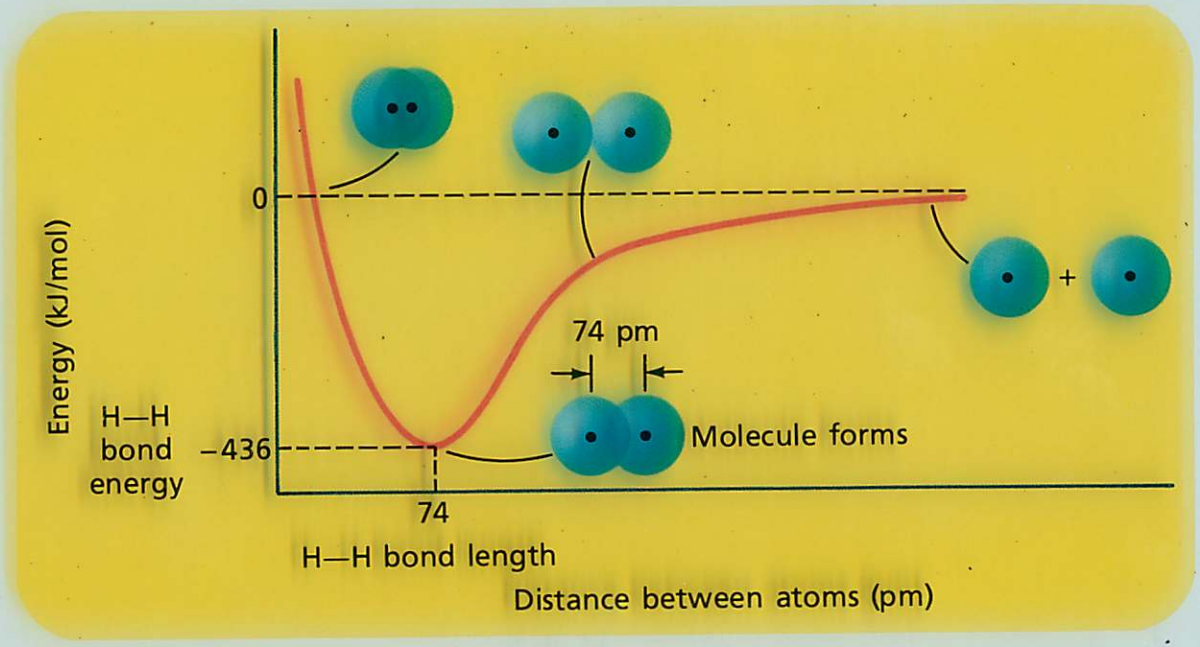


Figure 6-4

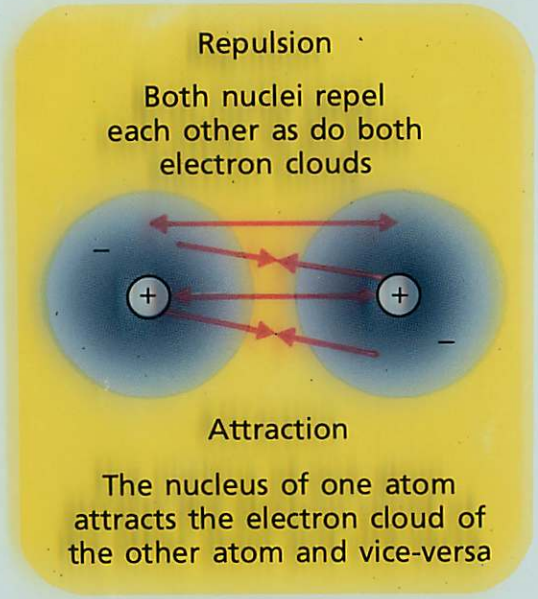


Figure 6-5

Bond Length → Avg dist b/tween 2 bonded atoms

↳ Varies w/ compounds (why? - Space orbital fills) (E-neg)

Bond energy → energy required to Break a chemical Bond.

↳ given as KJ/mol.

to break an H-H bond it take 436 KJ/mol

How much does it take to break for 1 molecule?

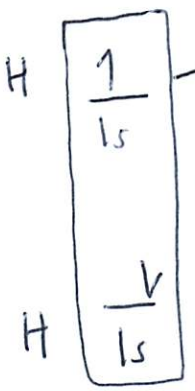
BE Longer as Bond Length shorter

What happens → Look @ orbital notation to understand

H $\frac{1}{1s}$ →

H $\frac{1}{1s}$

atoms



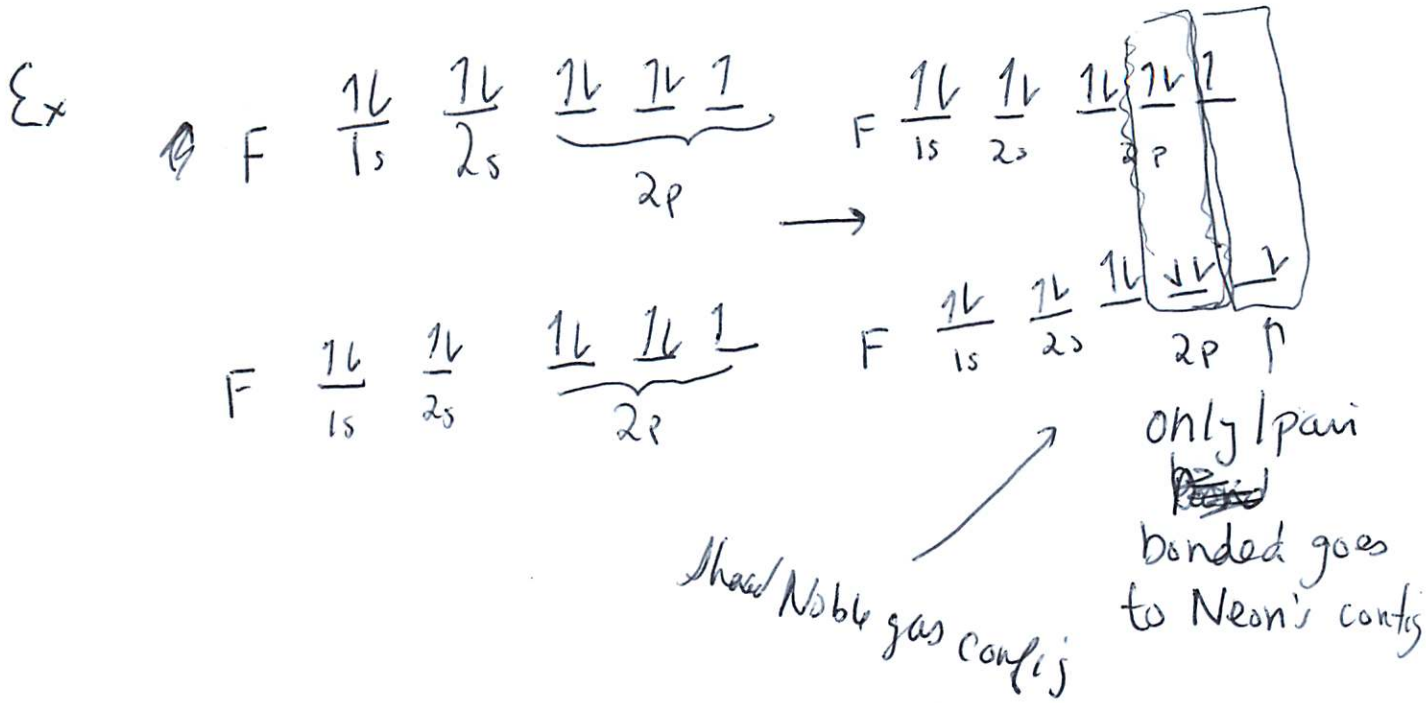
Box indicates unpaired e⁻ are now shared. Both occupy and fill each e⁻'s 1s orbital

B/c e⁻ are shared both 1s orbitals are filled and each H atom has stable $1s^2$ config (He)

the same happens for the other diatomic atoms

N, O, F, Cl, Br, I, At

Ex



These follow something called the Octet Rule

Octet Rule → compounds form so that each atom by either gaining losing or sharing an e^- has an octet of e^- in its highest E level

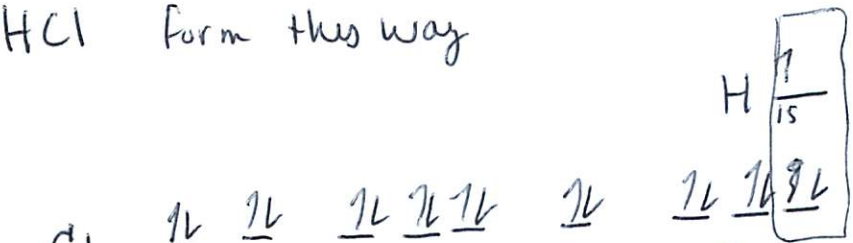
Exceptions BF_3 ← only 3 pairs

PCl_5 SF_5 → b/c d orbitals

↳ There are exceptions though

Discuss Later

HCl Form this way



Shared but polar b/c Cl e-neg

Lewis Structures

↳ e⁻ dot for molecules

H· H:H :Cl: :Cl:Cl: :Cl–Cl:

atom molecule atom molecule

Bond shown as line

H—H

Lewis Structure H—H :Cl—Cl: ← unshared e⁻ that belongs to Cl atom only

Shared e⁻ are represented as dash (bond)

Structural formula no dots for e⁻ only show bonds

Cl–Cl

Drawing Lewis Structures

Draw Iodomethane CH_3I (C is central)

1. How many elements present, which ones

1 C, 3 H, 1 I

2 Draw e⁻ dot for each

·C· H· :I:

skip →

3 Determin # of valence e^- to be combined

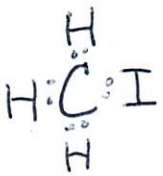
C $1 \times 4e^- = 4e^-$

I $1 \times 7e^- = 7e^-$

H $3 \times 1e^- = 3e^-$

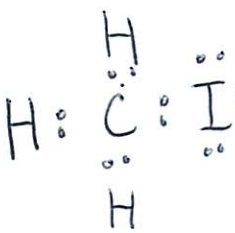
$14e^-$

4 Arrange atoms to form skeleton xture w/ those e^- pos
element central (C always central)
connect atoms w/ e^- dot pairs

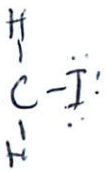


$8e^-$ so far

5 Add unshared pairs of e^- so each nonmetal atom (not H) has $8e^-$



$14e^-$ or



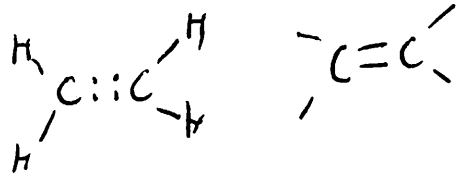
↑
Do a count to check

Ex Lewis xture for NH_3, ICl, SiF_4

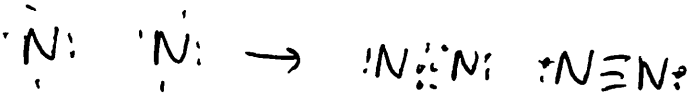
Rules of thumb → N 3 bonds, 1 unshared pair (unless (+) 4 bonds)
O 2 " 2 " " unles (−) → 1 bond)
H only 1 bond no free e^-
C always central, Always 4 bonds

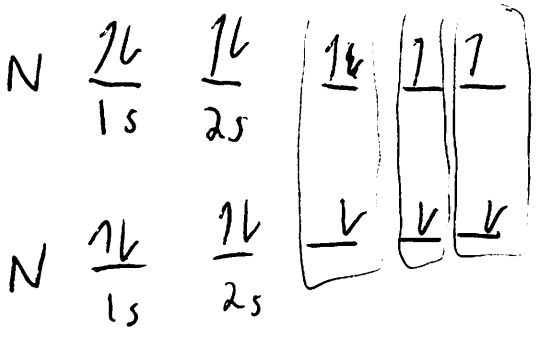
Multiple Bonds

Double and Triple bonds

↳ 2 pairs shared Ex Ethene 

Triple N can form triple bonds (3 pairs shared) Take more energy to break multiple bonds





C forms triple bonds too

H-C≡C-H ethyne

multiple bonds are possible

if get too many valence e⁻ when drawing Lewis structure

Do multiple bonds

Multiple formalin antiseptic CH_2O

Lewis Dot ·C· H· :O:

#valence

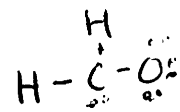
C 1 × 4e⁻

H 2 × 1e⁻

O 1 × 6e⁻

12e⁻

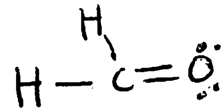
skeleton H HC:O

add valence 

count 14e⁻ too many? Double Bond

subtract 1 lone pair until get correct #.

Move another pair b/tween C or O, C−N to get multiple bond



−12e⁻

Ex CO_2, HCN

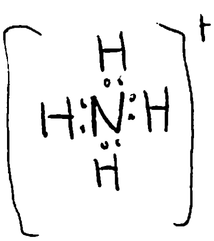
Polyatomic Ions − charged group of covalently bonded e⁻

Poly to draw Lewis Xture of. subtract (if posion) or add e (if anion) when we count e⁻

NH_4^+

·N: H·

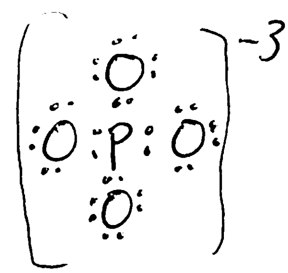
5 × 1 } 9e⁻
1 × 4 −1e⁻
8e⁻



PO_4^{-3}

·P: :O:

5 × 1 } 29e⁻
6 × 4 } +3e⁻
32e⁻



pg 174 Q 4, 5

In solids, these Ionic compounds form crystal lattices

see fig 6-10. B/c charges are like they are

The attraction of these charges in solid give them very high, MT, BP.

$NaCl \rightarrow$ BP 801
BP 1413
Sucrose 186°C

Na melts
Cl melts
NaCl melts

Ionic forces are very strong - Hold them packed tightly together

Dissolve in water easily b/c Polarity of H_2O

make good conductors

Pg 179 Q2

~~Give Criss Cross Weight~~

~~↳ Write out charges through~~

~~Show Criss Cross Method.~~

3rd Type of Bonding Metallic Bonding

↳ Metals → great conductors, malleable

↳ suggest mobility of e^-

Have vacant p + d orbitals @ their disposal

easily give up e^-, (low IE), do not attract e^- (low Eneg)

Valence e^- are given up and shared. They are free to move about and do not stay localized to one atom.

Called an e^- sea

15 NaCl Ion Packing and Crystal Diagram

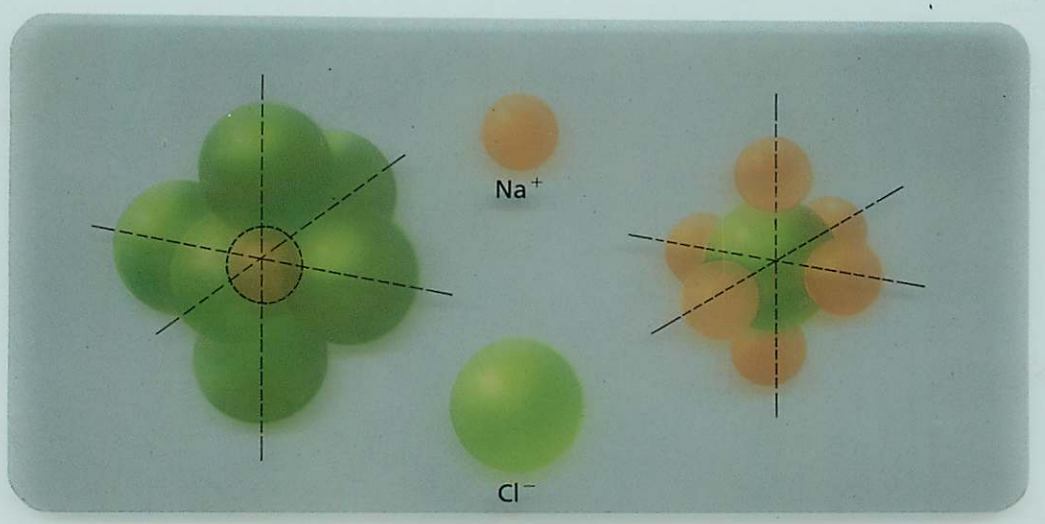


Figure 6-9

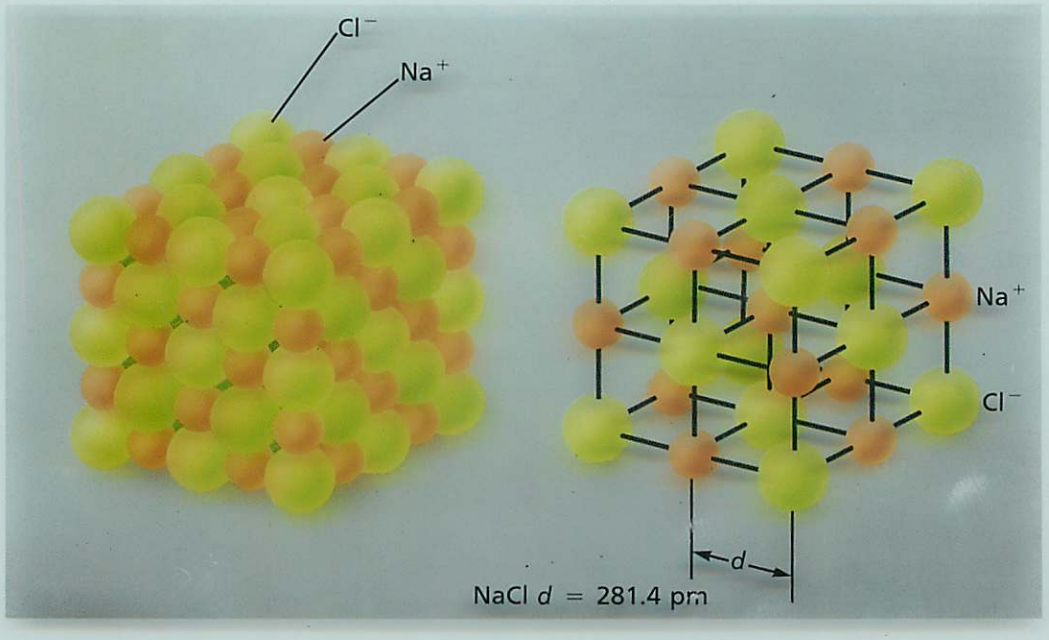
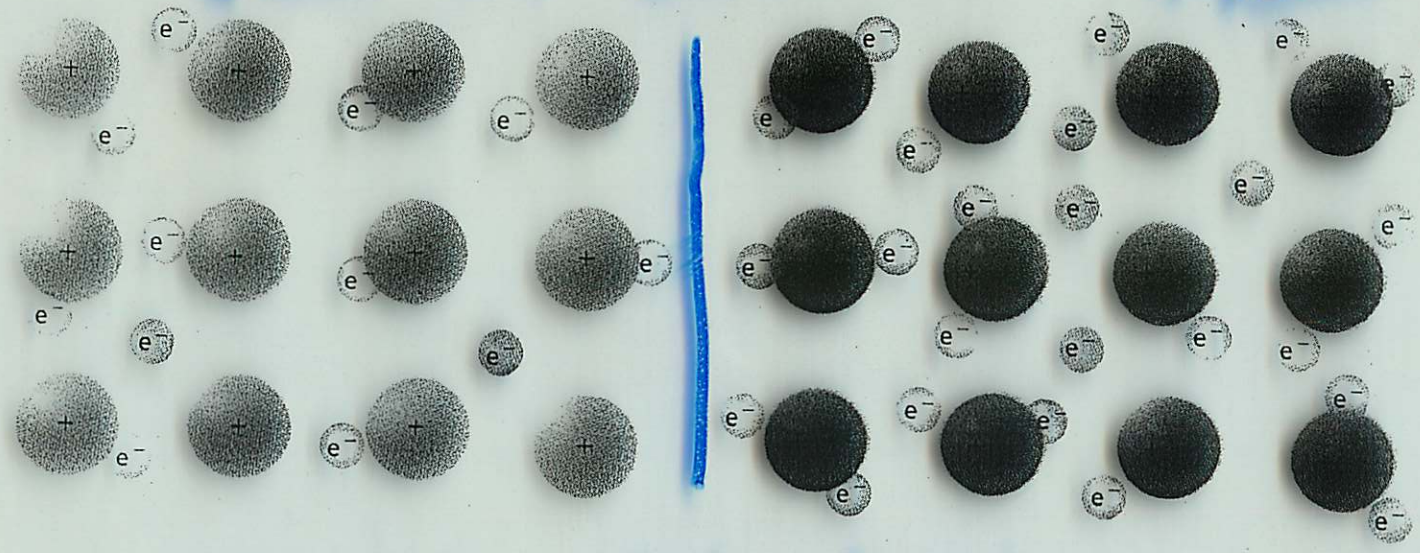


Figure 6-10

"Sea of electrons" fig 6.14



e⁻ are <u>delocalized</u>

they do not belong to any 1 atom

Strength of ~~Change~~ bond depends e⁻ shared + #of valence e⁻

Bond strength measured by how much heat it takes to vaporize a mole of that metal (kJ/mol)

↳ effect this - charge + # of valence e⁻

e⁻ sea explains high conductivity e⁻ carry charge
" heat conduction e⁻ move fast carry in heat

Malleability + Ductility
↳ B/c the bonds are not brittle
they can slide past one another w/o breaking bonds

181 Q 2+3

Columns for USER Table

molec type | unshare e⁻ pairs on central atom | Geom Type | Bond ∠ | Ex molec | Lewis structure | 3-D structure

Properties of compounds depend on 2 things

bonding -> ionic, covalent, etc

geometry - use 2 models which both fit different situations equally well

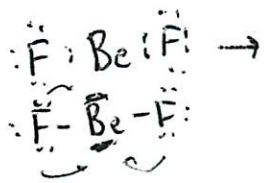
1 VSEPR - valence shell e⁻ pair repulsion

↳ geometry of compounds based on electrostatic repulsion - Bonds angle themselves in lowest energy possible (as far apart as possible)

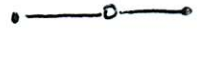
In compounds w/ 2 atoms

H_2 easy → H:H, H–H linear (A_2)

many varieties in ones w/ 3 or more atoms (chains)

BeF_2  → 0 lone e⁻

Type of mol: AB_2

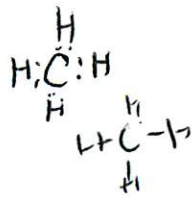
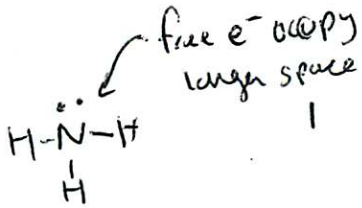
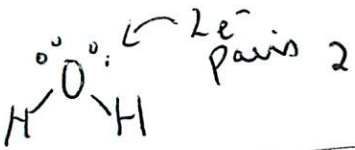
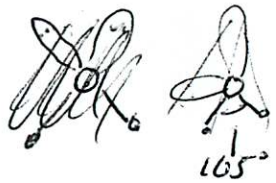
 linear 180°

equal repulsion

↓

BF_3  0 AB_3

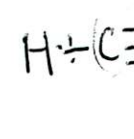
 triangular planar 120°

		lone pair	Type	Bond angle adopted	
CH_4		0	AB_4	109.5	 tetrahedral
NH_3	 free e⁻ occupy larger space	1	AB_3E	107°	 ~~trigonal~~ pyramidal
H_2O	 2 e⁻ pairs	2	AB_2E_2	105°	 Bent

to use table 6-4 it is necessary to know Lewis structure classify it to # of bonds around central atom

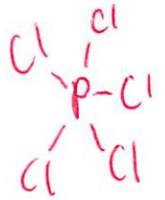
what type of bond does HCN + H_2CO have

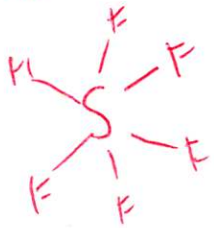
HCN



2 0 → linear

3 0 → triangular planar

PCl_5  AB_5 trigonal Pyramidal 

SF_6  AB_6 Octahedral 90°

Inter molecular forces - forces of attraction between molecules
↳ weaker forces than those that bond atoms in molecules

Strogest IM force → between polar molecules

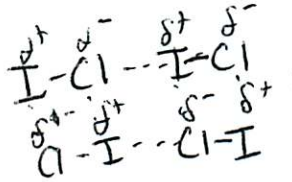
Dipole - Equal but opposite charges separated by a short distance

Like HCl

unequal distribution of e^-

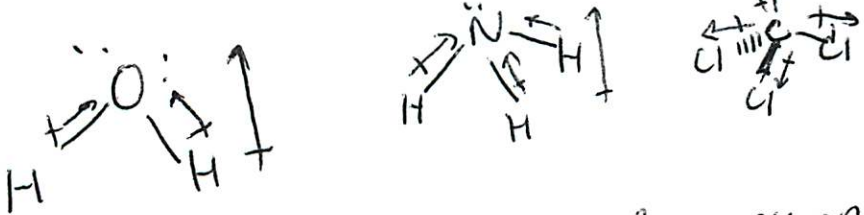
1. → Dipole-Dipole forces - attractive forces between polar molecules

Compare $\overset{\text{BP } 97^\circ C}{\overset{+}{I}-Cl}$ to $Br-Br$ $\quad -34.6^\circ C \quad 59^\circ C \quad Cl-Cl$



forces hold them together

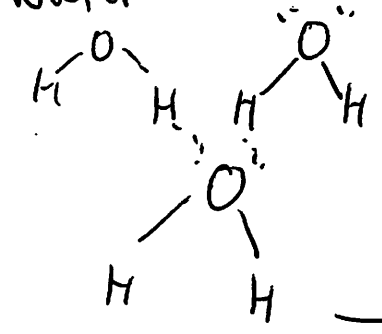
polarity in molecule - goes to more e-neg atom



↳ can induce dipoles in nonpolar molecules

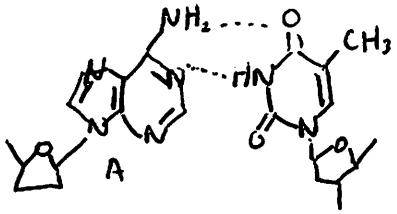
2. Hydrogen bonding - IM force attraction between an H bonded to a strong e⁻ negaton + unshared pair of e⁻ on other strong enegaton

Water rep by dashed line

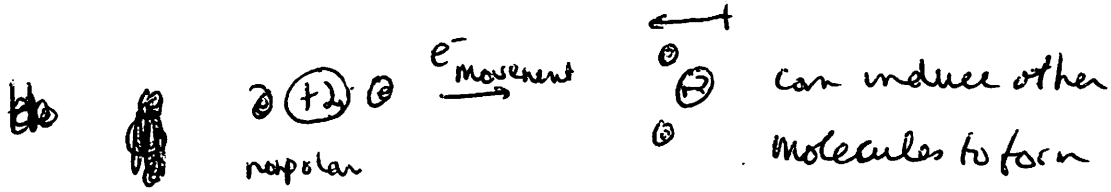


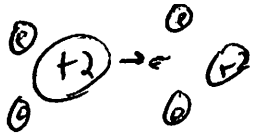
→ give water its high BP, Density etc surface tension

→ DNA



3. London dispersion forces → (always exp) constant motion of e⁻ and creation of instantaneous dipoles and induced dipoles





London higher in molecules w/ more e⁻

BP He −269

Ar −186 more e⁻ more LDF